Name : KEY Date: ________ Period: ____

What is an Isotope?

protons: determine which element an atom is - determines the atomic number
electrons: determine the charge on an atom

Protons = electrons - atom is neutral
Protons> electrons - atom is positively charged (lost electrons)
Protons< electrons - atom is negatively charged (gained electrons)

neutrons - determines the atomic mass of an atom. Determines the isotope number of the atom.
Isotope: an atom of an element with a certain number of neutrons.

NOTE: All atoms of an element are isotopes of that element.

Most elements have 1, 2 or 3 naturally occurring isotopes. This means that in any sample of the element these naturally occurring isotopes are all present, typically in the same % ratio.

$\ast$ (p has a mass of 1 amu
n has a mass of 1 amu) $\ast$

For example:
The element carbon has three isotopes: amu=atomic mass unit

${}^{12}_{6}C$ = 6 neutrons, mass number is 12, is called Carbon -12 isotope 90% abundance

${}^{13}_{6}C$ = 7 neutrons, mass number is 13, is called carbon-13 isotope 9% abundance

${}^{14}_{6}C$ = 8 neutrons, mass number is 14, is called carbon-14 isotope 1% abundance

% abundance means that in a sample of carbon (like a lump of coal or a diamond) 90% of the carbon atoms will be carbon-12, 9% will be carbon-13 and 1% will be carbon-14. Since not all the atoms in a sample of an element have the same mass, we have to calculate an average atomic mass for the element. The average atomic mass is calculated taking into account the different percents of each isotope present.

$$\text{Avg. Atomic mass} = \underbrace{\frac{\%\text{ of isotope \#1}}{100}}_{\text{dec form}} \times \overbrace{(\text{mass isotope \#1})}^{\text{mass \#}} + \frac{\%\text{ of isotope \#2}}{100} \times (\text{mass Isotope \#2}) + \frac{\%\text{ of isotope \#3}}{100} \times (\text{mass Isotope \#3})$$

An example of this type of calculation is provided on the next page.

$$(.90 \times 12) + (.09 \times 13) + (.01 \times 14) = 12.11 \text{ amu}$$

The weighted average atomic mass is calculated by multiplying the decimal equivalent of each isotope times its mass and adding up all the results for all the naturally occurring isotopes.

AMU = atomic mass unit = mass of one proton = mass of one neutron

Example:

A sample of Cesium, Cs, has the following % abundance:

Cs-132 = 20.0%; Cs-133 = 75.3%; Cs-134 = 4.7%.

Calculate the weighted average atomic mass.

Weighted average atomic mass =

(0.20 x 132) + (0.753 x 133) + (0.047 x 134) = 132.85 amu

Determine the average atomic mass for the following mixtures of isotopes:

1.) Au - 197 = 50%, Au 198 = 50%

49.98% 50.02%

197 x .50 = 98.5
198 x .50 = 99.0
197.5 amu

2.) Fe - 55 = 15%, Fe - 56 = 85%

55 x .15 = 8.25
56 x .85 = 47.6
55.85 amu

atomic mass unit

3.) H-1 = 99%, H-2- = 0.8%, H-3 = 0.2%

.008 .002

1 x .99 = .99
2 x .008 = .016
3 x .002 = .006
1.012 amu

4.) N-14 = 95%, N-15 = 3%, N-16 = 2%

14 x .95 = 13.3
15 x .03 = .45
16 x .02 = .32
14.07 amu

5.) C-12 = 98%, C-14 = 2%

12 x .98 = 11.76
14 x .02 = .28
12.04 amu

Do not worry about sig figs on these problems!